Chapter-3

Lesson-1

I – Chemical Equations

In a chemical reaction, atoms are neither created nor destroyed. Thus, the total mass of the products will equal the total mass of the reactants. **This is the law of conservation of mass**.

A) Coefficients and Subscripts:

Coefficients are the numbers in front of a formula in an equation. The coefficients tell how many of each species are present. ***Subscripts*** are the numbers that follow each element in a formula and tell how many atoms of the element are present in the formula.

The coefficient 3 indicates three molecules of carbon dioxide.

$$3\ CO_2$$

Subscripts 1(implied) and 2 indicate one carbon atom and two oxygen atoms.

B) Balancing Equations:

An equation is balanced by determining the coefficients that provide equal numbers of each type of atom on each side of the equation. Subscripts should **never** be changed when balancing equations.

- Rusting of iron in air:

  $\underline{\ 4\ }\ Fe + \underline{\ 3\ }\ O_2 \rightarrow \underline{\ 2\ }\ Fe_2O_3$

- Thermal decomposition of potassium chlorate:

  $\underline{\ 2\ }\ KClO_3 \rightarrow \underline{\ 2\ }\ KCl + \underline{\ 3\ }\ O_2$

- Complete combustion of propane:

  $\underline{\ 1\ }\ C_3H_8 + \underline{\ 5\ }\ O_2 \rightarrow \underline{\ 3\ }\ CO_2 + \underline{\ 4\ }\ H_2O$

C) States of Reactants and Products:

The species in a reaction can be in the solid, liquid, gaseous, or aqueous state. The symbols (s), (l), (g), and (aq) are used to indicate the state of each species in the equation.

- Neutralization of vinegar with baking soda:

$\underline{\ 1\ }\ HC_2H_3O_{2(aq)} + \underline{\ 1\ }\ NaHCO_{3(s)} \rightarrow$

$\underline{\ 1\ }\ NaC_2H_3O_{2(aq)} + \underline{\ 1\ }\ H_2O_{(l)} + \underline{\ 1\ }\ CO_{2(g)}$

- Reaction of zinc metal and nitric acid:

$\underline{\ 1\ }\ Zn_{(s)} + \underline{\ 2\ }\ HNO_{3(aq)} \rightarrow \underline{\ 1\ }\ Zn(NO_3)_{2(aq)} + \underline{\ 1\ }\ H_{2(g)}$

II – Patterns of Chemical Reactions

In order to predict the products of a chemical reaction in which the reactants are given, the reaction type must be recognized. Simple patterns of chemical reactivity exist and can be used to determine the reaction type. Combination reactions, decomposition reactions, and combustion reactions of hydrocarbons and alcohols are three common reaction types.

A) Combination Reactions:

Two or more reactants combine to form a single product.

$$\mathbf{A + B \rightarrow C}$$

- Rusting of iron in air:

$$\mathbf{4\ Fe_{(s)} + 3\ O_{2(g)} \rightarrow 2\ Fe_2O_{3(s)}}$$

- Burning of graphite in air:

$$\mathbf{C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}}$$

- Wetting of lime:

$$\mathbf{CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(s)}}$$

- Mixing of ammonia and hydrogen chloride fumes:

$$\mathbf{HCl_{(g)} + NH_{3(g)} \rightarrow NH_4Cl_{(s)}}$$

B) Decomposition Reactions:

A single reaction breaks apart (often when heated).

$$\mathbf{C \rightarrow A + B}$$

- Heating of potassium chlorate:

$$\mathbf{2\ KClO_{3(s)} \rightarrow 2\ KCl_{(s)} + 3\ O_{2(g)}}$$

- Heating of limestone:

$$\mathbf{CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}}$$

- Heating of baking soda:

$$\mathbf{2\ NaHCO_{3(s)} \rightarrow Na_2O_{(s)} + H_2O_{(g)} + 2\ CO_{2(g)}}$$

- Heating of smelling salts:

$$\mathbf{(NH_4)_2CO_{3(s)} \rightarrow NH_{3(g)} + H_2O_{(g)} + 2\ CO_{2(g)}}$$

C) Combustion Reactions:

A hydrocarbon or alcohol reacts with oxygen to form carbon dioxide and water.

$$\mathbf{C_xH_y + O_2 \rightarrow CO_{2(g)} + H_2O_{(g)}}$$

- Combustion of propane:

$$\mathbf{C_3H_{8(g)} + 5\ O_{2(g)} \rightarrow 3\ CO_{2(g)} + 4\ H_2O_{(g)}}$$

- Combustion of rubbing alcohol:

$$\mathbf{2\ CH_3CH(OH)CH_{3(l)} + 9\ O_{2(g)} \rightarrow 6\ CO_{2(g)} + 8\ H_2O_{(g)}}$$

The reactions above are called ***complete combustion*.**

Complete combustion occurs when **only** CO_2 and H_2O are produced. Complete combustion occurs when the supply of oxygen is plentiful as in air.

However, when the supply of oxygen is limited, then ***incomplete combustion*** occurs in which some CO and/or some C is produced. Water is always a product of complete or incomplete combustion.

- Incomplete combustion of propane:

$$\mathbf{C_3H_{8(g)} + 4\ O_{2(g)} \rightarrow CO_{2(g)} + 2\ CO_{(g)} + 4\ H_2O_{(g)}}$$

III – Formula Weights and Percent Weights

A) Formula Weight:

- The ***formula weight*** (**FW**) of a substance is the sum of the atomic weights of each atom in its chemical formula.
- If the chemical formula is that of a molecule, then the formula weight is also called the ***molecular weight*** (**MW**).

B) Percent By Mass from a Compound's Formula:

The percent by mass (percent composition) of each element in a compound can be determined from its formula.

$$\textbf{\%element} = \frac{\textbf{(\# of atoms)(atomic weight)}}{\textbf{(formula weight of compound)}}$$

Q1: What is the percent composition by mass of HClO?

A1: FW = 1.01 + 35.45 + 16.00 = 52.46

$$\%H = \frac{1.01}{52.46} \times 100 = \boxed{1.93\%\ H}$$

$$\%Cl = \frac{35.45}{52.46} \times 100 = \boxed{67.58\%\ Cl}$$

$$\%O = \frac{16.00}{52.46} \times 100 = \boxed{30.50\%\ O}$$

Q2: Calculate the percent by mass of each element in KNO_3?

A2: FW = 39.10 + 14.01 + 3×16.00 = 101.11

$$\%K = \frac{39.10}{101.11} \times 100 = \boxed{38.67\%\ K}$$

$$\%N = \frac{14.01}{101.11} \times 100 = \boxed{13.86\%\ N}$$

$$\%O = \frac{48.00}{101.11} \times 100 = \boxed{47.47\%\ O}$$

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IV – Avogadro's Number and the Mole

A) Avogadro's Number:

Q1: What do 12.01 g C, 32.06 g S, & 63.55 g Cu have in common?

A1: They all contain the same number of atoms!

- This number of atoms is called ***Avogadro's number***. The number of atoms present in a sample of an element is equal to Avogadro's number when the mass of the sample is equal to the element's gram atomic mass!
- Avogadro's number = 6.022×10^{23} per mol [6.022×10^{23} mol^{-1}]
- The concept of Avogadro's number was known long before the value of the number was known.
- Several remarkable experiments were performed in order to find the value of Avogadro's number.
  - The Electroplating Experiment - To electroplate 1.000 g of solid silver from a solution of a silver salt (**$Ag^{+}_{(aq)} + e^{-} \rightarrow Ag^{0}_{(s)}$**) a certain amount of current (**I**) is required for a certain amount of time (**t**). Since I × t = q (**q = charge in coulombs**), and since a single electron has a charge of 1.602×10^{-19} coulombs, then the total number of electrons needed to plate a gram Ag can be found. Since $1e^{-}$ is needed to plate one Ag atom, then a simple ratio can be set up to calculate Avogadro's number of Ag atoms.
  - The X-ray Diffraction Experiment - Using x-ray diffraction, it is possible to determine the pattern of and the distances between the atoms in a crystal. If the volume of a certain number of atoms can be determined in cm^3, then using density and atomic mass Avogadro's number of atoms can be found.

B) The Mole:

- 1 mole = 6.022×10^{23} = Avogadro's number
  - ***1 mole C atoms = 6.022×10^{23} atoms = 12.011 g***
  - ***1 mole H atoms = 6.022×10^{23} atoms = 1.008 g***
  - ***1 mole O atoms = 6.022×10^{23} atoms = 15.999 g***
- The mass of 1 mole of an element is called the element's ***molar mass*** (**MM or** M).
  - ***The molar mass of carbon is 12.011 g/mol***
  - ***The molar mass of hydrogen is 1.008 g/mol***
  - ***The molar mass of oxygen is 15.999 g/mol***
- The mass of 1 mole of a compound is called the compound's ***molar mass*** (**MM or** M).
  - ***The molar mass of $C_6H_{12}O_6$ is 180.156 g/mol***
- Molar mass is a conversion unit (# g / 1 mole)
- [molar mass]$^{-1}$ is a conversion unit (1 mole / # g)

Q2: What is the mass in grams of 0.746 mole $C_6H_{12}O_6$?

A2: $0.746 \text{ mole} \times \dfrac{180.156 \text{ g}}{1 \text{ mole}} = 134 \text{ g } C_6H_{12}O_6$

V - Empirical and Molecular Formulas

Q3: What is the empirical and molecular formula of an unknown compound whose composition is 85.60% C and 14.40% H by mass and whose molecular mass is 84.18 g/mol?

A3: $85.60 \text{ g C} \times \dfrac{1 \text{ mole}}{12.01 \text{ g}} = 7.127 \text{ mol} \div 7.127 \text{ mol} = 1.000 \rightarrow 1$

$\boxed{\mathbf{CH_2}}$

$14.40 \text{ g H} \times \dfrac{1 \text{ mole}}{1.008 \text{ g}} = 14.26 \text{ mol} \div 7.127 \text{ mol} = 2.001 \rightarrow 2$

CH_2 has a molar mass of 14.03 g/mol

The unknown compound's molecular mass is 6 times larger.

Thus the molecular formula of the unknown compound is $\boxed{\mathbf{C_6H_{12}.}}$

Q4: What is the empirical formula of a compound whose composition is 81.68% C and 18.32% H by mass?

A4: $81.68 \text{ g C} \times \frac{1 \text{ mole}}{12.01 \text{ g}} = 6.801 \text{ mol} \div 6.801 \text{ mol} = 1.000$

$18.32 \text{ g H} \times \frac{1 \text{ mole}}{1.008 \text{ g}} = 18.16 \text{ mol} \div 6.801 \text{ mol} = 2.\underline{\underline{670}}$

$\boxed{\mathbf{C_?H_?}}$

The decimal 0.670 is approximately ⅔. The **denominator** of the approximate fraction is multiplied by the previous answers.

$1.000 \times 3 = 3$

$2.\underline{\underline{670}} \times 3 \approx 8$

$\boxed{\mathbf{C_3H_8}}$

Q5: What is the empirical formula of the **hydrate** washing soda ($Na_2CO_3 \bullet xH_2O$) if a 2.558 g sample is heated (*but not too hot … why?*) to a constant weight of 0.948 g?

A5: If a hydrate is heated to a constant weight, the water will be driven off and only the salt will remain.

$0.948 \text{ g } Na_2CO_3 \times \frac{1 \text{ mole}}{106 \text{ g}} = 0.00894 \text{ mol} \div 0.00894 \text{ mol} = 1$

$1.610 \text{ g } H_2O \times \frac{1 \text{ mole}}{18.02 \text{ g}} = 0.0893 \text{ mol} \div 0.00894 \text{ mol} \approx 10$

$\boxed{\mathbf{Na_2CO_3 \bullet 10H_2O}}$

Q6: What is the empirical formula of a compound that is composed of 1.481 g N and 2.538 g O?

A6: $1.481 \text{ g N} \times \frac{1 \text{ mole}}{14.01 \text{ g}} = 0.1057 \text{ mol} \div 0.1057 \text{ mol} = 1.000 \times 2 = 2$

$2.538 \text{ g O} \times \frac{1 \text{ mole}}{16.00 \text{ g}} = 0.1586 \text{ mol} \div 0.1057 \text{ mol} = 1.\underline{\underline{500}} \times 2 = 3$

$\boxed{\mathbf{N_2O_3}}$

Q7: What is the empirical formula of a 27.80 g compound that is composed of 4.14 g of phosphorus and chlorine?

A7: The mass of chlorine can be found using the law of conservation of mass:

27.80 g compound - 4.14 g phosphorus = **23.66 g chlorine**

$$4.14 \text{ g P} \times \frac{1 \text{ mole}}{30.97 \text{ g}} = 0.134 \text{ mol} \div 0.134 \text{ mol} = 1.00$$

$$23.66 \text{ g Cl} \times \frac{1 \text{ mole}}{35.45 \text{ g}} = 0.667 \text{ mol} \div 0.134 \text{ mol} = 4.981 \approx 5$$

$\boxed{\mathbf{PCl_5}}$

Q8: What is the empirical formula of the **hydrate** $AuCl_3 \bullet xH_2O$ if a 15.000 g sample is heated to a constant weight of 13.407 g?

A8: $$13.407 \text{ g } AuCl_3 \times \frac{1 \text{ mole}}{303.32 \text{ g}} = 0.044201 \text{ mol} \div 0.044201 \text{ mol} = 1$$

$$1.593 \text{ g } H_2O \times \frac{1 \text{ mole}}{18.02 \text{ g}} = 0.08840 \text{ mol} \div 0.044201 \text{ mol} = 2$$

$\boxed{\mathbf{AuCl_3 \bullet 2H_2O}}$

Q9: What are the empirical and molecular formulas of a compound that is composed of 30.450% N and 69.550 % O by mass and whose molecular mass is 46.010 g/mol?

A9: $$30.450 \text{ g N} \times \frac{1 \text{ mole}}{14.01 \text{ g}} = 2.1734 \text{ mol} \div 2.1734 \text{ mol} = 1.000$$

$$69.550 \text{ g O} \times \frac{1 \text{ mole}}{16.00 \text{ g}} = 4.3469 \text{ mol} \div 2.1734 \text{ mol} = 2.000$$

$\boxed{\textbf{Empirical} = \mathbf{NO_2} \qquad \textbf{Molecular formula} = \mathbf{N_2O_4}}$

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VI – Determining Formulas by Combustion Analysis

The empirical and molecular formulas of many organic compounds ($C_xH_yO_z$ or C_xH_y) can be determined through ***combustion analysis***.

Q1: 1.000 g of an organic acid burns to give 1.466 g of CO_2 and 0.6001 g H_2O. What is the empirical formula of the acid?

A1: $C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$ ← ***unbalanced skeleton equation***

All the C in the CO_2 & all the H in the H_2O started in the mystery acid!

$$1.466\text{ g } CO_2 \times \frac{1\text{ mol } CO_2}{44.01\text{ g } CO_2} \times \frac{1\text{ mol C}}{1\text{ mol } CO_2} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = \boxed{\mathbf{0.4001\text{ g C}}}$$

$$0.6001\text{g } H_2O \times \frac{1\text{ mol } H_2O}{18.02\text{ g } H_2O} \times \frac{2\text{ mol H}}{1\text{ mol } H_2O} \times \frac{1.01\text{ g H}}{1\text{ mol H}} = \boxed{\mathbf{0.0673\text{ g H}}}$$

The mass of oxygen is found using the Law of Conservation of Mass.

$$\text{mass of oxygen} = 1.000\text{ g} - [0.4001 + 0.0674\text{ g}] = \boxed{\mathbf{0.533\text{ g O}}}$$

The empirical formula can now be determined.

$$0.4001\text{ g C} \times \frac{1\text{ mole}}{12.01\text{ g}} = 0.03331\text{ mol} \div 0.0333\text{ mol} = 1.00$$

$$0.0673\text{ g H} \times \frac{1\text{ mole}}{1.01\text{ g}} = 0.0666\text{ mol} \div 0.0333\text{ mol} = 2.00 \quad \boxed{\mathbf{CH_2O}}$$

$$0.533\text{ g O} \times \frac{1\text{ mole}}{16.00\text{ g}} = 0.0333\text{ mol} \div 0.0333\text{ mol} = 1.00$$

☞ ***To determine the molecular formula of a compound from its empirical formula, the molecular mass of the compound __must__ be known.***

<u>Q2</u>: What is the molecular formula for the organic acid whose empirical formula was determined to be CH_2O and whose molecular mass is 60.0 g/mol?

<u>A2</u>: If the empirical formula's mass is 30.0 g/mol, and the molecular formula's mass is 60.0 g/mol (***double***), then the molecular formula is ... $\boxed{\mathbf{C_2H_4O_2}}$

<u>Q3</u>: 0.476 g of a hydrocarbon burns to give some CO_2 and 0.428 g of H_2O. What are the empirical and molecular formulas of the hydrocarbon if its molar mass is 120 g/mol?

<u>A3</u>: $C_xH_y + O_2 \rightarrow CO_2 + H_2O$ ← ***<u>unbalanced skeleton equation</u>***

<u>All the H in the H_2O started in the mystery hydrocarbon</u>!

$$0.428 \text{ g } H_2O \times \frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O} \times \frac{2 \text{ mol H}}{1 \text{ mol } H_2O} \times \frac{1.01 \text{ g H}}{1 \text{ mol H}} = \boxed{\mathbf{0.0480\text{ g H}}}$$

<u>The mass of carbon is found using the Law of Conservation of Mass</u>.

$$\text{mass of carbon} = 0.476 \text{ g} - 0.0480 \text{ g} = \boxed{\mathbf{0.428\text{ g C}}}$$

<u>The empirical formula can now be determined</u>.

$$0.428 \text{ g C} \times \frac{1 \text{ mole}}{12.01 \text{ g}} = 0.0356 \text{ mol} \div 0.0356 \text{ mol} = 1.00 \times 3 = 3$$

$$0.0480 \text{ g H} \times \frac{1 \text{ mole}}{1.01 \text{ g}} = 0.0475 \text{ mol} \div 0.0356 \text{ mol} = 1.32 \times 3 = 4$$

$\boxed{\mathbf{C_3H_4}}$

<u>The molecular formula can now be determined</u>.

C_3H_4 has a mass of 40 g/mol

$\boxed{\mathbf{C_9H_{12}\text{ has a mass of 120 g/mol}}}$

Q4: In a combustion analysis of a 0.1665 g sample of aspartame, it was found that 0.3180 g of CO_2, 0.0840 g of H_2O, and 0.0290 g of N_2 were produced. Determine the empirical and molecular formulas of aspartame, if its molar mass is 294 g/mol?

A4: $C_wH_xO_yN_z + O_2 \rightarrow CO_2 + H_2O + N_2 \leftarrow$ ***unbalanced skeleton equation***

***All the C in the CO_2 & all the H in the H_2O started in the aspartame*!**

$$0.3180\text{ g } CO_2 \times \frac{1\text{ mol } CO_2}{44.01\text{ g } CO_2} \times \frac{1\text{ mol C}}{1\text{ mol } CO_2} \times \frac{12.01\text{ g C}}{1\text{ mol C}} = \boxed{\mathbf{0.08678\text{ g C}}}$$

$$0.0840\text{ g } H_2O \times \frac{1\text{ mol } H_2O}{18.02\text{ g } H_2O} \times \frac{2\text{ mol H}}{1\text{ mol } H_2O} \times \frac{1.01\text{ g H}}{1\text{ mol H}} = \boxed{\mathbf{0.00942\text{ g H}}}$$

***The mass of oxygen is found using the Law of Conservation of Mass*.**

mass of oxygen = 0.1665 g – [0.08678 g + 0.00942 g + 0.0290 g]

$= \boxed{\mathbf{0.0413\text{ g O}}}$

***The empirical and molecular formulas can now be determined*.**

$$0.08678\text{ g C} \times \frac{1\text{ mole}}{12.01\text{ g}} = 0.007226\text{ mol} \div 0.00104\text{ mol} = 6.95 \approx 7$$

$$0.00942\text{ g H} \times \frac{1\text{ mole}}{1.01\text{ g}} = 0.00933\text{ mol} \div 0.00104\text{ mol} = 8.97 \approx 9$$

$$0.0413\text{ g O} \times \frac{1\text{ mole}}{16.00\text{ g}} = 0.00258\text{ mol} \div 0.00104\text{ mol} = 2.48 \approx 2.5$$

$$0.0290\text{ g } N_2 \times \frac{1\text{ mole}}{28.02\text{ g}} = 0.00104\text{ mol} \div 0.00104\text{ mol} = 1.00 = 1$$

The empirical formula $\mathbf{C_{14}H_{18}O_5N_2}$ has a mass of 294 g/mol

The molecular formula of aspartame is $\mathbf{C_{14}H_{18}O_5N_2}$.

Q5: A 2.612 g sample of a solid copper oxide - $Cu_xO_{y(s)}$, when heated in a stream of hydrogen gas, yielded solid copper and 0.592 g of water. What was the empirical formula of the copper oxide?

A5: $Cu_xO_{y(s)} + H_{2(g)} \rightarrow Cu_{(s)} + H_2O_{(g)} \leftarrow$ ***unbalanced skeleton equation***

***All the O in the H_2O started in the copper oxide*!**

$$0.592\,g\,H_2O \times \frac{1\,mol\,H_2O}{18.02\,g\,H_2O} \times \frac{1\,mol\,O}{1\,mol\,H_2O} \times \frac{16.00\,g\,O}{1\,mol\,O} = \boxed{\mathbf{0.527\,g\,O}}$$

***The mass of copper is found using the Law of Conservation of Mass*.**

$$\text{mass of copper} = 2.612\text{ g} - 0.527\text{ g} = \boxed{\mathbf{2.085\text{ g Cu}}}$$

***The empirical formula can now be determined*.**

$$2.085\text{ g Cu} \times \frac{1\text{ mole}}{63.55\text{ g}} = 0.03281\text{ mol} \div 0.03281\text{ mol} = 1.000$$

$$0.527\text{ g O} \times \frac{1\text{ mole}}{16.00\text{ g}} = 0.0329\text{ mol} \div 0.03281\text{ mol} = 1.00$$

The empirical formula is $\boxed{\mathbf{CuO}.}$

Q6: What is the IUPAC name for the copper oxide formula determined in question-5?

A6: copper(II) oxide

Q7: What is the traditional name for the copper oxide formula determined in question-5?

A7: cupric oxide

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VII – Calculations Based on Chemical Equations

A) Stoichiometric Relations in Reactions:

Q1: How many moles of CO_2 are produced when 1.65 mole of C_3H_8 [propane] burns in air?

A1: $C_3H_{8(g)} + 5\ O_{2(g)} \rightarrow 3\ CO_{2(g)} + 4\ H_2O_{(l)}$

$$1.65 \text{ mol } C_3H_8 \times \frac{\mathbf{3\ mol\ CO_2}}{\mathbf{1\ mol\ C_3H_8}} = \boxed{4.95 \text{ mol } CO_2}$$

Q2: How many moles of oxygen gas are produced when 27.5 g $KClO_3$ is heated and decomposes?

A2: $2\ KClO_{3(s)} \rightarrow 2\ KCl_{(s)} + 3\ O_{2(g)}$

$$27.5 \text{ g } KClO_3 \times \frac{1 \text{ mole}}{122.6 \text{ g}} \times \frac{\mathbf{3\ mole\ O_2}}{\mathbf{2\ mol\ KClO_3}} = \boxed{0.336 \text{ mol } O_2}$$

Q3: At STP, how many liters of CO_2 would be released if 454 g $NaHCO_3$ [baking soda] reacts with excess $HC_2H_3O_2$ [vinegar]?

A3: $HC_2H_3O_{2(aq)} + NaHCO_{3(s)} \rightarrow NaC_2H_3O_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$

$$454 \text{ g } NaHCO_3 \times \frac{1 \text{ mol}}{84.0 \text{ g}} \times \frac{\mathbf{1\ mol\ CO_2}}{\mathbf{1\ mol\ NaHCO_3}} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = \boxed{121 \text{ L } CO_2}$$

Q4: How many grams of CaO [lime] would be needed to neutralize 95000.0 L of 0.00100 M H_2SO_4 ["acid rain"]? **FYI… M = *mol/L***

A4: $CaO_{(s)} + H_2SO_{4(aq)} \rightarrow CaSO_{4(s)} + H_2O$

$$95000 \text{ L} \times 0.00100 \frac{\text{mol}}{1\text{L}} H_2SO_4 \times \frac{\mathbf{1\ mol\ CaO}}{\mathbf{1\ mol\ H_2SO_4}} \times \frac{56.1 \text{ g}}{1 \text{ mol}} = \boxed{5330 \text{ g}}$$

Q5: How many grams of solid Zn would be needed to reduce 27.0 g Ag from a concentrated $AgNO_3$ solution?

A5: $Zn_{(s)} + 2\ AgNO_{3(aq)} \rightarrow Zn(NO_3)_{2(s)} + 2\ Ag$

$$27.0\text{ g Ag} \times \frac{1\text{ mol}}{108\text{ g}} \times \frac{\mathbf{1\ mol\ Zn}}{\mathbf{2\ mol\ Ag}} \times \frac{65.4\text{ g}}{1\text{ mol}} = \boxed{8.18\text{ g Zn}}$$

B) The Limiting Reactant and the Theoretical Yield:

- Ordinarily, reactants are not present in the exact ratio required for a reaction. Instead, one reactant is in excess; some of it is left when the reaction is over. The other reactant, called the ***limiting reactant***, is completely converted to give the ***theoretical yield*** of product.
- The ***theoretical yield*** of a product in a reaction is stoichiometrically determined from the ***limiting reactant***.
- To calculate which reactant is the ***limiting reactant***, determine which reactant will "run out" first during the reaction.

Q6: If 25.0 g Ag reacts with 25.0 g I_2 according to produce $AgI_{(s)}$, then
(a) determine the limiting reactant and
(b) calculate the theoretical yield of AgI
(c) determine how much original reactant remains unreacted

A6: $\underline{2}\ Ag_{(s)} + \underline{1}\ I_{2(s)} \rightarrow 2\ AgI_{(s)}$

(a) Calculate the number of moles of both reactants …

$$25.0\text{ g Ag} \times \frac{1\text{ mole}}{107.9\text{ g}} = 0.232\text{ mole Ag} \div \underline{2}\text{ mole} = 0.116$$

coefficients

$$25.0\text{ g } I_2 \times \frac{1\text{ mole}}{253.8\text{ g}} = 0.0985\text{ mole } I_2 \div \underline{1}\text{ mole} = 0.0985$$

I_2 is the limiting reactant.

(b) $$0.0985\text{ mole } I_2 \times \frac{\mathbf{2\ mole\ AgI}}{\mathbf{1\ mole\ I_2}} \times \frac{234.8\text{ g}}{1\text{ mole}} = \boxed{46.3\text{ g AgI}}$$

(c) $$0.0985\text{ mole } I_2 \times \frac{\mathbf{2\ mole\ Ag}}{\mathbf{1\ mole\ I_2}} \times \frac{107.9\text{ g}}{1\text{ mole}} = 21.3\text{ g Ag used up}$$

$\boxed{3.7\text{ g Ag left}}$

Q7: 0.0403 g Mg reacts with 10.0 mL of 3.00 M $HCl_{(aq)}$.

$$Mg + 2\,HCl \rightarrow MgCl_2 + H_2$$

(a) Determine which reactant is the limiting reagent.
(b) Calculate (at STP) the theoretical yield (***in mL***) of the gaseous product.

A7: (a) $10.0\text{ ml HCl} \times \dfrac{1\text{L}}{10^3\text{ mL}} \times \dfrac{3.00\text{ mol}}{1\text{ L}} = 0.0300\text{ mol} \div 2\text{ mol} = 0.0150$

Mg is the limiting reactant.

$$0.0403\text{ g Mg} \times \frac{1\text{ mole}}{24.3\text{ g}} = 0.00166\text{ mol} \div 1\text{ mol} = 0.00166$$

(b) $0.00166\text{ mol} \times \dfrac{\mathbf{1\ mol\ H_2}}{\mathbf{1\ mol\ Mg}} \times \dfrac{22.4\text{ L}}{1\text{ mol}} = 0.0372\text{ L} \rightarrow \boxed{37.2\text{ mL}}$

Q8: 45.0 g C_2H_6 & 48.0 g O_2 are mixed and then ignited. The reaction that occurs is $2\,C_2H_6 + 7\,O_2 \rightarrow 4\,CO + 6\,H_2O$. (incomplete combustion)

(a) Determine which reactant is the limiting reagent.
(b) Explain the clue to the identity of the limiting reagent.
(c) Calculate of theoretical yield (***in grams***) of water.

A8: (a) $45.0\text{ g } C_2H_6 \times \dfrac{1\text{ mol}}{30.0\text{ g}} \div 2\text{ moles} = 0.750$

O_2 is the limiting reactant!

$$48.0\text{ g } O_2 \times \frac{1\text{ mol}}{32.0\text{ g}} \div 7\text{ moles} = 0.214$$

(b) The fact that carbon monoxide is a product instead of carbon dioxide indicates that O_2 is the limiting reagent!

(c) $48.0\text{ g } O_2 \times \dfrac{1\text{ mol}}{32.0\text{ g}} \times \dfrac{\mathbf{6\ mol\ H_2O}}{\mathbf{7\ mol\ O_2}} \times \dfrac{18.0\text{ g}}{1\text{ mol}} = \boxed{23.1\text{ g } H_2O}$

Q9: 0.250 g $ZnCl_2$ is added to a 15.0 mL $AgNO_3$ (0.150 M) solution.

(a) Write the balanced equation for the reaction.
(b) Determine which reactant is the limiting reagent.
(c) Calculate the theoretical yield (***in grams***) of the precipitate.

A9: (a) $ZnCl_{2(aq)} + 2\ AgNO_{3(aq)} \rightarrow Zn(NO_3)_{2(aq)} + 2\ AgCl_{(s)}$

(b) $0.250\text{ g } ZnCl_2 \times \frac{1\text{ mol } ZnCl_2}{136.3\text{ g}} = 0.00183\text{ mol} \div 1\text{ mol} = 0.00183$

$AgNO_3$ limits the reaction!

$0.0150\text{ L} \times \frac{0.150\text{ mol } AgNO_3}{1\text{ L}} = 0.00225\text{ mol} \div 2\text{ mol} = 0.00113$

(c) $0.00225\text{ mol} \times \frac{\mathbf{2\ mol\ AgCl}}{\mathbf{2\ mol\ AgNO_3}} \times \frac{143.4\text{ g}}{1\text{ mol}} = \boxed{0.323\text{ g AgCl}}$

C) The Actual Yield and the Percent Yield:

- The ***actual yield*** may be less than the ***theoretical yield***.

$$\textbf{\% Yield} = \frac{\textbf{Actual Yield}}{\textbf{Theoretical Yield}} \times \textbf{100}$$

Q10: Suppose the actual yield of H_2 in question-7 was 30.0 mL. What was the percent yield?

A10: $\%\text{ Yield} = \frac{30.0\text{ mL } H_2}{37.2\text{ mL } H_2} \times 100 = \boxed{80.6\%}$

Q11: Suppose the actual yield of H_2O in question-8 was 20.0 g. What was the percent yield?

A11: $\%\text{ Yield} = \frac{20.0\text{ g } H_2O}{23.1\text{ g } H_2O} \times 100 = \boxed{86.6\%}$

Q12: Suppose the % yield of Ag in question-5 was 77.0%. How many grams of Zn would be needed to still produce 27.0 g Ag?

A12: $77.0\% = \frac{27.0\text{ g Ag}}{X\text{ g Ag}} \times 100 \rightarrow X = \boxed{35.1\text{ g of theoretical Ag}}$

$35.1\text{ g Ag} \times \frac{1\text{ mol}}{108\text{ g}} \times \frac{\mathbf{1\ mol\ Zn}}{\mathbf{2\ mol\ Ag}} \times \frac{65.4\text{ g}}{1\text{ mol}} = \boxed{10.6\text{ g Zn needed}}$